Topic 6: The rate and extent of chemical change

AQA GCSE Chemistry (8462), topic 4.6

The rate and extent of chemical change

- calculate mean rates of reaction from tables and graphs, and draw tangents to find instantaneous rates (higher tier);
- use **collision theory** to explain how **concentration**, **pressure**, surface area, temperature and catalysts change the rate;
- plan and interpret the required practical on concentration and rate, by two methods;

The rate and extent of chemical change (2)

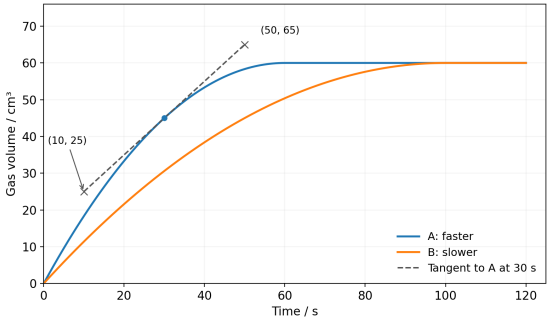
- describe reversible reactions and **dynamic equilibrium**;
- (higher tier) predict how changing conditions moves the position of an equilibrium.

Measuring the rate of a reaction

- The **rate of reaction** tells you how fast reactants are used up or products are made.
- The **mean rate** over a period of time is:
- The quantity can be a mass (rate in g/s) or a gas volume (cm^3/s); at higher tier it can also be an amount in moles (mol/s).
- **Worked example 1.** The volume of gas rises from 12 cm^3 at 10 s to 36 cm^3 at 30 s.

$$\text{mean rate} = \frac{\text{quantity of reactant used or product formed}}{\text{time taken}}$$

Figure: Measuring the rate of a reaction

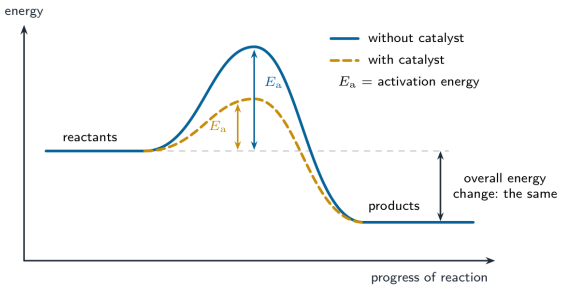


- Gas volume against time for three concentrations, with a **tangent** showing **instantaneous rate**.
- Read the axes and units first.

What changes the rate?

- **Collision theory** says that particles react only when they collide, and only if they collide with at least the **activation energy** .
- Anything that makes successful collisions happen more often makes the reaction faster.
- Temperature has two effects, and the second is the bigger one: more of the collisions have enough energy to react.
- **Concentration** does not change the **kinetic energy** of the particles; it only makes collisions more frequent.

Figure: What changes the rate?

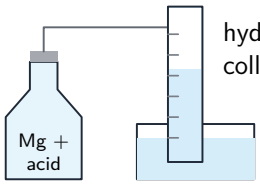


- Catalysed and uncatalysed reaction profiles — lower **activation energy**, same overall change.
- Match each part of the diagram to the text.

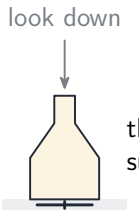
Required practical: concentration and rate

- Before you start, write a **hypothesis** , for example: "If the acid is more concentrated, the reaction will be faster, because there will be more frequent collisions."
- **Method 1: measuring a gas.** Add a 3 cm length of cleaned **magnesium ribbon** to 50 cm^3 of **dilute hydrochloric acid** in a conical flask.
- Quickly fit a bung with a **delivery tube** that leads into an upside-down **measuring cylinder** full of water (or use a **gas syringe**).
- Record the volume of hydrogen every 10 seconds until it stops changing.

Figure: Required practical: concentration and rate



gas-volume method: read the volume every 10 s until constant



turbidity method: time until the cross disappears

- Collect the gas and read the volume.
- Or time until the cross disappears.

Reversible reactions

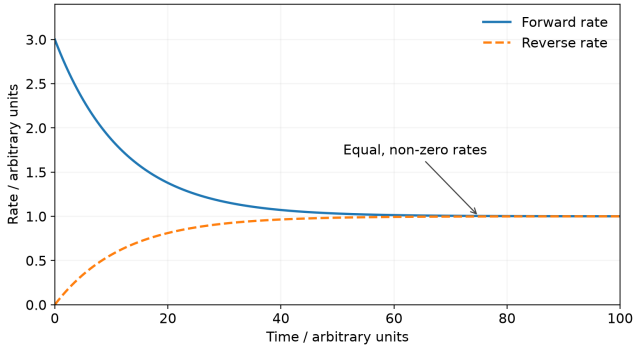
- In a **reversible reaction** , the products can react to form the original reactants again.
- The symbol $\rightleftharpoons$ shows this:
- Heating **ammonium chloride** breaks it down into two gases; when they cool, they join to form **ammonium chloride** again.
- If a reaction is exothermic in one direction, it is endothermic in the other, and the same amount of energy is transferred. **Hydrated copper sulfate** is blue.

$$\text{ammonium chloride} \rightleftharpoons \text{ammonia} + \text{hydrogen chloride}$$

Dynamic equilibrium

- In a **closed system** , nothing can get in or out.
- When a **reversible reaction** happens in a **closed system**, it eventually reaches **dynamic equilibrium** : the forward and reverse reactions go on at exactly the same rate.
- The amounts of reactants and products then stay constant, although they are usually not equal.
- At the start, only reactants are present, so the forward reaction is fast.

Figure: Dynamic equilibrium



- Equal continuing forward and reverse rates in a **closed system**.
- Match each part of the diagram to the text.

Changing an equilibrium (higher tier)

- There are 4 gas molecules on the left and 2 on the right, so high **pressure** gives more ammonia.
- Low temperature gives a higher **yield** of ammonia, but then the reaction is very slow. Industry uses a compromise of about 450 °C, with an iron **catalyst** to speed it up.
- A catalyst does not change the **equilibrium position**; it only helps the system reach equilibrium sooner.

Sources for this lesson

- AQA, *GCSE Chemistry (8462) specification*, section 4.6, and the required practicals handbook (Required practical 5).
- Minnesota Historical Society, "Washburn A Mill explosion, 2 May 1878", and Mill City Museum, Minneapolis.
- US Chemical Safety Board, *Combustible Dust Hazard Study* (2006).
- H. L. Le Chatelier, "Sur un énoncé général des lois des équilibres chimiques", *Comptes rendus* 99 (1884).

Common errors

- Saying that a higher temperature only makes particles collide more often; more collisions also have enough energy to react.
- Saying that a **catalyst** gives more product; it only makes the reaction faster.
- Joining two points on the curve instead of drawing a **tangent** that touches it at one point.

Common errors (2)

- Saying that reactions stop at equilibrium; both reactions continue at the same rate.
- Saying that the amounts of reactants and products are equal at equilibrium; they are constant, not equal.
- Counting all molecules, including solids, when predicting the effect of **pressure**; count gas molecules only.

Key terms

- **Rate of reaction**: how fast a reactant is used up or a product is formed.
- **Collision theory**: particles react only when they collide with enough energy.
- **Activation energy**: the minimum energy that colliding particles need to react.

Key terms (2)

- **Catalyst**: a substance that speeds up a reaction without being used up.
- **Reversible reaction**: a reaction in which the products can react to form the reactants.
- **Dynamic equilibrium**: the state of a **reversible reaction** in a **closed system** when the forward and reverse rates are equal.
- **Le Chatelier principle**: a system at equilibrium responds to oppose any change in conditions.

Check yourself

- 1. 0.60 g of magnesium is used up in 40 s.
- What is the **mean rate**? 2.
- Explain, using **collision theory**, why powdered marble reacts faster with acid than marble chips. 3.
- Give two reasons why raising the temperature increases the rate. 4.